Chem 173: Precipitation of Ionic Compounds

To predict if a precipitate will form when 2 solutions are mixed:

- Calculate the concentration of individual ions in the combined solution (remember to divide mol of ion by the \textit{TOTAL} solution volume)
- Calculate the ion product, Q (same form as $K_{sp}$, but ion concentrations here are not necessarily equilibrium ion concentrations)
- Compare Q to $K_{sp}$
- If Q > $K_{sp}$, the solution is \textit{supersaturated}, ion concentrations are greater than equilibrium concentrations, reaction will proceed in reverse to reach equilibrium, \textit{precipitation will occur}.
- If Q < $K_{sp}$, the solution is \textit{unsaturated}, ion concentrations are less than equilibrium concentrations, reaction will proceed forward to reach equilibrium, \textit{more solid will dissolve}.
- If Q = $K_{sp}$, the solution is \textit{saturated}, the solution is at equilibrium, ion concentrations are equilibrium concentrations, \textit{no more solid will dissolve or precipitate}.

examples:

- Will a precipitate form when 150 mL of 0.10 M Pb(NO$_3$)$_2$ are combined with 100 mL of 0.20 M NaCl?
  \[ \text{PbCl}_2 (s) \rightleftharpoons \text{Pb}^{2+} (aq) + 2 \text{Cl}^– (aq); \quad K_{sp} = 1.2 \times 10^{-5} \]
  mol Pb$^{2+}$ (aq) = (0.150 L sol'n)(0.10 mol Pb$^{2+}$/L sol'n) = 0.015 mol Pb$^{2+}$
  \[ [\text{Pb}^{2+}] = 0.015 \text{mol}/0.250 \text{L} = 0.060 \text{M} \]
  mol Cl$^–$ (aq) = 0.100 L sol'n)(0.20 mol Cl$^–$/L sol'n) = 0.020 mol Cl$^–$
  \[ [\text{Cl}^–] = 0.020 \text{mol}/0.250 \text{L} = 0.080 \text{M} \]
  \[ Q = [\text{Pb}^{2+}][\text{Cl}^–]^2 = (0.060)(0.080)^2 = 3.8 \times 10^{-4} \]
  Q > $K_{sp}$; PbCl$_2$ will precipitate from solution

- What minimum concentration of Na$_2$CO$_3$ (aq) is required to cause precipitation of BaCO$_3$ (s) from a solution of 1.0 x $10^{-5}$ M BaCl$_2$ (aq)?
  \[ \text{BaCO}_3 (s) \rightleftharpoons \text{Ba}^{2+} (aq) + \text{CO}_3^{2–} (aq); \quad K_{sp} = 2.6 \times 10^{-9} \]
  logic: At the point where Q = $K_{sp}$ the solution is at equilibrium. Calculate the concentration of CO$_3^{2–}$ that will satisfy this relationship:
  \[ Q = [\text{Ba}^{2+}][\text{CO}_3^{2–}] = 1.0 \times 10^{-5}[\text{CO}_3^{2–}] \]
  \[ [\text{CO}_3^{2–}] = 2.6 \times 10^{-4} \text{M} \]
  This is the concentration of Na$_2$CO$_3$ that will result in a saturated solution (solid and ions in equilibrium). So . . . any [Na$_2$CO$_3$] greater than 2.6 x $10^{-4}$ M will result in Q > $K_{sp}$ and precipitation of BaCO$_3$.

- What compound, if any, will precipitate when 80.0 mL of 1.0 x $10^{-5}$ M Ba(OH)$_2$ (aq) is added to 20.0 mL 1.0 x $10^{-5}$ M Fe$_2$(SO$_4$)$_3$?
  \[ \text{BaSO}_4 (s) \rightleftharpoons \text{Ba}^{2+} (aq) + \text{SO}_4^{2–} (aq); \quad K_{sp} = 1.1 \times 10^{-10} \]
  \[ \text{Fe(OH)}_3 (s) \rightleftharpoons \text{Fe}^{3+} (aq) + 3 \text{OH}^– (aq); \quad K_{sp} = 2.6 \times 10^{-39} \]
  mol Ba$^{2+}$ = 0.0800 L sol'n(1.0 x $10^{-5}$ mol Ba$^{2+}$/L sol'n) = 8.0 x $10^{-7}$ mol Ba$^{2+}$
  \[ [\text{Ba}^{2+}] = 8.0 \times 10^{-7} \text{mol}/0.100 \text{L} = 8.0 \times 10^{-6} \text{M} \]
  mol OH$^–$ = 0.0800 L sol'n(1.0 x $10^{-5}$ mol Ba(OH)$_2$/L sol'n)(2 mol OH$^–$/1 mol Ba(OH)$_2$) = 1.6 x $10^{-6}$ mol OH$^–$
  \[ [\text{OH}^–] = 1.6 \times 10^{-6} \text{mol}/0.100 \text{L} = 1.6 \times 10^{-5} \text{M} \]
  mol Fe$^{3+}$ = 0.0200 L sol'n(1.0 x $10^{-5}$ mol Fe$_2$(SO$_4$)$_3$/L sol'n)(2 mol Fe$^{3+}$/1 mol Fe$_2$(SO$_4$)$_3$) = 4.0 x $10^{-7}$ mol Fe$^{3+}$
  \[ [\text{Fe}^{3+}] = 4.0 \times 10^{-7} \text{mol}/0.100 \text{L} = 4.0 \times 10^{-6} \text{M} \]
  mol SO$_4^{2–}$ = 0.0200 L sol'n(1.0 x $10^{-5}$ mol Fe$_2$(SO$_4$)$_3$/L sol'n)(3 mol SO$_4^{2–}$/1 mol Fe$_2$(SO$_4$)$_3$) = 6.0 x $10^{-7}$ mol SO$_4^{2–}$
  \[ [\text{SO}_4^{2–}] = 6.0 \times 10^{-7} \text{mol}/0.100 \text{L} = 6.0 \times 10^{-6} \text{M} \]
  For BaSO$_4$: \[ Q = [\text{Ba}^{2+}][\text{SO}_4^{2–}] = (8.0 \times 10^{-6})(6.0 \times 10^{-6}) = 4.8 \times 10^{-11} \]
  Q < $K_{sp}$; BaSO$_4$ \text{DOES NOT} precipitate from the solution
  For Fe(OH)$_3$: \[ Q = [\text{Fe}^{3+}][\text{OH}^–]^3 = (4.0 \times 10^{-6})(1.6 \times 10^{-5})^3 = 1.6 \times 10^{-20} \]
  Q > $K_{sp}$; Fe(OH)$_3$ \text{DOES} precipitate from the solution